Heat of Combustion of Low M.M. Fuels

OBJECTIVE: To apply your knowledge of calorimetry to analytically determining quantity of energy per mole of ethanol or methanol, ΔH_{comb}.

BACK GROUND:
Concerns about finite petroleum reserves have led to a search for renewable energy sources to supplement or replace traditional hydrocarbon fuels. Ethanol has been proposed as an alternative fuel source for automobiles because it can be prepared from fermentation of almost any plant product. In some areas, "gasohol" consisting of 90% gasoline and 10% ethanol is already in use.
The heat of combustion ΔH_{comb} for a fuel is defined as enthalpy change for the following reaction when balances:

$$\text{Fuel} + O_{2\,(g)} \rightarrow CO_{2\,(g)} + H_2O_{(l)}$$

Safety: Fuels are very volatile and flammable use caution.

PROCEDURES:
1. Fill a calorimeter with 150ml of ice cold DI water (no ice cube though). Cover the calorimeter with the lid and get the initial temp of the water by leaving the thermometer in the water bath. Place the calorimeter on the wire gauze on the iron ring.

2. Measure out 3-5ml of Ethanol or methanol, which ever you are assigned. And add it to your evaporating dish.

3. Place the evaporating dish under the calorimeter (about 5-10cm).

4. Light the fuel in your evaporating dish with a wooden splint and the wrap the sides with Al foil (see Mr. Golden for example).

5. Let the evaporating dish burn out completely and record the highest temp the water reaches.

6. Wait for five minutes and repeat steps 1-5 two more time.

DATA:

	Trial 1	**Trail 2**	**Trial 3**
Vol. H_2O mass	ml g	ml g	ml g
Vol. Fuel (CH_3OH or C_2H_5OH)	ml	ml	ml
T_i H_2O	°C	°C	°C
T_f H_2O	°C	°C	°C
ΔT H_2O	°C	°C	°C
q_{water}	J	J	J
mol_{fuel}	mol	mol	mol
ΔH_{comb}	kJ/mol	kJ/mol	kJ/mol

CALCULATIONS:

We will assume that our calorimeters absorb no heat in order to simplify calculations.

$$\Delta H_{comb} = \frac{q_{fuel}\ (kJ)}{mol_{fuel}}$$

To get q_{fuel}:

We can then use the "q" of water to find the joules of energy released from the fuel.

$$q_{H_2O} = C_p \times m \times \Delta T$$

Then we can convert the joules to kilojoules (kJ)

mol_{fuel}:

Next we need mols of fuel we will assume the fuel is at 25°C and at that temp the density of ethanol is 0.789g/ml while methanol is 0.792g/ml.

Multiple the vol. of fuel by its density, then divide by molar mass of fuel.

$\Delta H_{combustion}$

Finally divide the kJ of energy, q_{fuel}, by the moles of fuel used. That will give you $\Delta H_{combustion}$

Now determine the ΔH_{comb} for each of your trials and find the average $\Delta H_{combustion}$ for your experiment.

QUESTIONS:

1. How does your experimental value of ΔH_{comb} compared to the actual value of your fuel; CH_3OH 726kJ/mol or C_2H_5OH 1368kJ/mol?

2. Why do you suppose methanol has a lower $\Delta H_{comb.}$ than ethanol?

3. What are three sources of error in your experiment?

4. If you burned 500ml of ethanol, what amount of energy (kJ) would be produced? (use the density, molar mass, and $\Delta H_{comb.}$ of ethanol)

5. Briefly define the 1^{st} law of thermodynamics in you own words.

6. If a 450g piece of cobalt at a temperature of 205°C was placed in a water bath containing 250g of water at 30°C? What is losing heat, the water or the cobalt?
   Specific heat of Co = 0.42 J/g°C
   Specific heat of water = 4.184 J/g°C

7. What would the final temperature be of the cobalt and water?